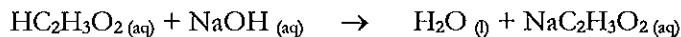


## CHM129

## Acid-Base Equilibrium

## Weak Acid – Strong Base and Weak Base – Strong Acid Titrations

1. Consider the titration of 40.0 mL of 0.100M acetic acid ( $K_a = 1.8 \times 10^{-5}$ ) with a 0.200 M NaOH solution.



Determine the pH at the following points:

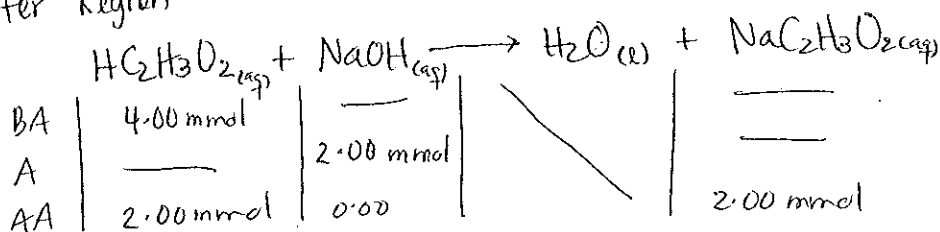
- Initial pH (no NaOH added)
- pH after the addition of 10.00 mL of NaOH
- pH after the addition of 20.00 mL of NaOH
- pH after the addition of 30.00 mL of NaOH

Equivalence Point:  $C_a V_a = C_b V_b \Rightarrow V_b = \frac{C_a V_a}{C_b} = \frac{(0.100\text{M})(40.00\text{mL})}{(0.200\text{M})} = 20.0\text{mL NaOH}$

(a) Initial pH

$$[\text{H}_3\text{O}^+] = \sqrt{K_a [\text{acid}]} = \sqrt{(1.8 \times 10^{-5})(0.100)} = 1.3 \times 10^{-3}\text{M} \quad \text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(1.3 \times 10^{-3}) = \underline{2.87}$$

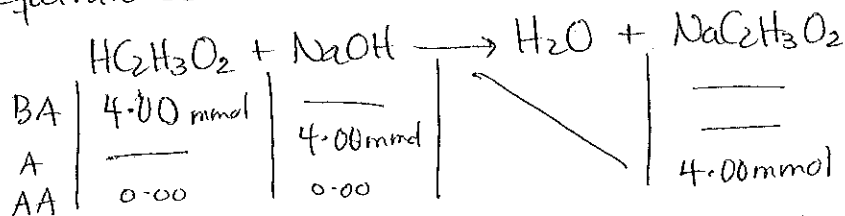
(b) Buffer Region



\* Note: Midpoint  
 $V_b = \frac{1}{2} V_{\text{eq}} = 10.00\text{mL}$   
 So,  $\text{pH} = \text{p}K_a$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]} = -\log(1.8 \times 10^{-5}) + \log \frac{(2.00\text{mmol})}{(2.00\text{mmol})} = \underline{4.74}$$

(c) Equivalence Point



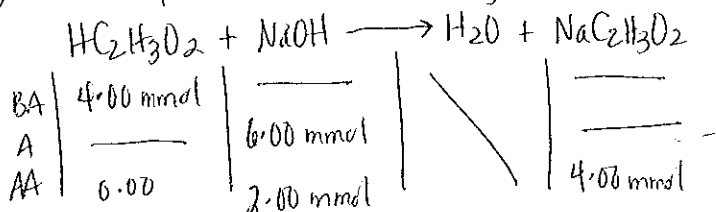
$$[\text{NaC}_2\text{H}_3\text{O}_2] = \frac{4.00\text{mmol}}{60.00\text{mL}} = 0.0667\text{M}$$

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

$$[\text{OH}^-] = \sqrt{K_b [\text{base}]} = \sqrt{(5.6 \times 10^{-10})(0.0667)} = 6.1 \times 10^{-6}\text{M}$$

$$\text{pOH} = -\log(6.1 \times 10^{-6}) = 5.22 \quad \Rightarrow \text{pH} = 14.00 - 5.22 = \underline{8.78}$$

(d) After Equivalence Point (Strong base excess)

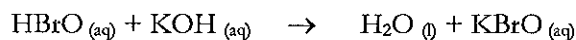


$$[\text{OH}^-] = [\text{NaOH}] = \frac{2.00\text{mmol}}{70.00\text{mL}} = 0.0286\text{M}$$

$$\text{pOH} = -\log(0.0286) = 1.544$$

$$\text{pH} = 14.00 - 1.544 = \underline{12.46}$$

2. Consider the titration of 35.0 mL of 0.140M hypobromous acid ( $K_a = 2.5 \times 10^{-9}$ ) with a 0.200 M KOH solution.



Determine the pH at the following points:

- Initial pH (no KOH added)
- pH after the addition of 13.00 mL of KOH
- pH after the addition of 24.50 mL of KOH
- pH after the addition of 35.00 mL of KOH

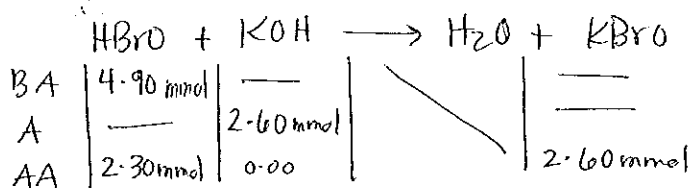
Equivalence Point

$$C_a V_a = C_b V_b \Rightarrow V_b = \frac{C_a V_a}{C_b} = \frac{(35.0 \text{ mL})(0.140 \text{ M})}{0.200 \text{ M}} = 24.50 \text{ mL}$$

(a) Initial pH

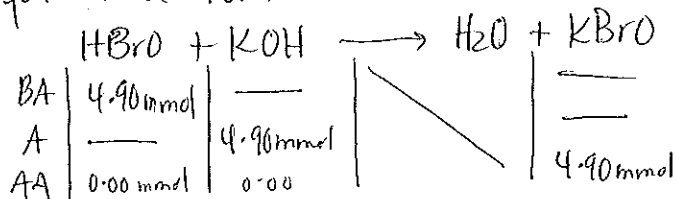
$$[\text{H}_3\text{O}^+] = \sqrt{K_a \times [\text{acid}]} = \sqrt{(0.140 \text{ M})(2.5 \times 10^{-9})} = 1.9 \times 10^{-5} \text{ M} \Rightarrow \text{pH} = -\log(1.9 \times 10^{-5}) = \underline{\underline{4.73}}$$

(b) Buffer Region



$$\text{pH} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]} = -\log(2.5 \times 10^{-9}) + \log \frac{(2.60)}{(2.30)} = \underline{\underline{8.66}}$$

(c) Equivalence Point



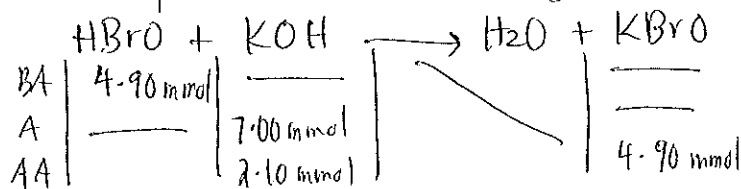
$$[\text{KBrO}] = \frac{4.90 \text{ mmol}}{59.50 \text{ mL}} = 0.0824 \text{ M}$$

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{2.5 \times 10^{-9}} = 4.0 \times 10^{-6}$$

$$[\text{OH}^-] = \sqrt{K_b \times [\text{KBrO}]} = \sqrt{(4.0 \times 10^{-6})(0.0824)} = 5.7 \times 10^{-4} \text{ M}$$

$$\text{pOH} = -\log(5.7 \times 10^{-4}) = 3.24 \Rightarrow \text{pH} = 14.00 - 3.24 = \underline{\underline{10.76}}$$

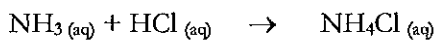
(d) After Equivalence Point (Strong base excess)



$$[\text{OH}^-] = [\text{KOH}] = \frac{2.10 \text{ mmol}}{70.00 \text{ mL}} = 0.0300 \text{ M}$$

$$\text{pOH} = -\log(0.0300) = 1.523 \Rightarrow \text{pH} = 14.00 - 1.523 = \underline{\underline{12.48}}$$

3. Consider the titration of 30.0 mL of 0.200M ammonia ( $K_b = 1.8 \times 10^{-5}$ ) with a 0.100 M HCl solution.



Determine the pH at the following points:

- Initial pH (no HCl added)
- pH after the addition of 30.00 mL of HCl
- pH after the addition of 60.00 mL of HCl
- pH after the addition of 70.00 mL of HCl

Equivalence Point:

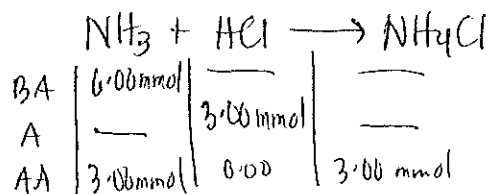
$$C_a V_a = C_b V_b \Rightarrow V_a = \frac{C_b V_b}{C_a} = \frac{(30.0 \text{ mL})(0.200 \text{ M})}{(0.100 \text{ M})} = 60.0 \text{ mL}$$

(a) Initial pH

$$[\text{OH}^-] = \sqrt{K_b \times [\text{base}]} = \sqrt{(1.8 \times 10^{-5})(0.200)} = 1.9 \times 10^{-3} \text{ M}$$

$$\text{pOH} = -\log(1.9 \times 10^{-3}) = 2.72 \Rightarrow \text{pH} = 14.00 - 2.72 = \underline{11.28}$$

(b) Buffer Region



$$K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}}$$

$$K_a = 5.6 \times 10^{-10}$$

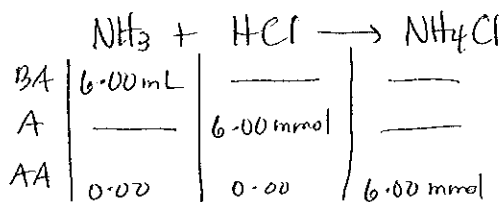
\* Note: Mid point

$$V_a = \frac{1}{2} V_{eq} = 30.0 \text{ mL}$$

$$\text{pH} = \text{p}K_a$$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]} = -\log(5.6 \times 10^{-10}) + \log \frac{(3.00)}{(3.00)} = \underline{9.25}$$

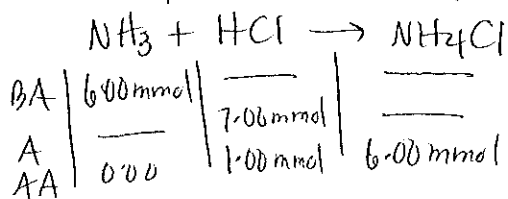
(c) Equivalence Point



$$[\text{NH}_4\text{Cl}] = \frac{6.00 \text{ mmol}}{90.00 \text{ mL}} = 0.0667 \text{ M}$$

$$[\text{H}_3\text{O}^+] = \sqrt{K_a \times [\text{acid}]} = \sqrt{(5.6 \times 10^{-10})(0.0667)} = 6.1 \times 10^{-6} \text{ M} \Rightarrow \text{pH} = -\log(6.1 \times 10^{-6}) = \underline{5.22}$$

(d) After Equivalence Point (Strong acid excess)



$$[\text{H}_3\text{O}^+] = [\text{HCl}] = \frac{1.00 \text{ mmol}}{100.00 \text{ mL}} = 0.0100 \text{ M}$$

$$\text{pH} = -\log(0.0100) = \underline{2.000}$$

